

ALL work must be shown to receive full credit. **Due at the beginning of lecture on Friday, November 2, 2001.**

PS11.1. For aqueous solutions of the following substances, write the dissociation reaction and indicate whether the substance behaves as an Arrhenius acid or base.

- a) $\text{HF}(aq) \rightleftharpoons \text{H}^+(aq) + \text{F}^-(aq)$ **acid**
b) $\text{HC}_6\text{H}_5\text{O}(aq) \rightleftharpoons \text{H}^+(aq) + \text{C}_6\text{H}_5\text{O}^-(aq)$ **acid**
c) $\text{Ba}(\text{OH})_2(aq) \rightarrow \text{Ba}^{2+}(aq) + 2\text{OH}^-(aq)$ **base**
d) $\text{LiOH}(aq) \rightarrow \text{Li}^+(aq) + \text{OH}^-(aq)$ **base**
e) $\text{H}_2\text{O}(aq) \rightleftharpoons \text{H}^+(aq) + \text{OH}^-(aq)$ **neutral**
f) $\text{H}_2\text{CO}_3(aq) \rightleftharpoons \text{H}^+(aq) + \text{HCO}_3^-(aq)$ **acid**

PS11.2. Calculate the pH and pOH in each of the following aqueous solutions. In each case, indicate whether the solution is acidic or basic.

- a) $[\text{H}^+] = 3.89 \times 10^{-5} \text{ M}$ **pH = 4.41**
pOH = 9.59
acidic
b) $[\text{OH}^-] = 8.34 \times 10^{-2} \text{ M}$ **pH = 12.92**
pOH = 1.08 **basic**
c) $[\text{OH}^-] = 1.50 \times 10^{-7} \text{ M}$ ($[\text{OH}^-]$ in milk) **pH = 7.18**
pOH = 6.82 **basic**
d) $[\text{H}^+] = 9.39 \times 10^{-10} \text{ M}$ **pH = 9.02**
pOH = 4.97 **basic**
e) $[\text{H}^+] = 4.0 \text{ M}$ **pH = -0.60**
pOH = 14.6 **acidic**
f) $[\text{OH}^-] = 10.1 \text{ M}$ **pH = 15.0**
pOH = -1.00 **basic**

PS11.3. Calculate the $[\text{H}^+]$ and $[\text{OH}^-]$ in each of the following aqueous solutions.

- pH = -log $[\text{H}^+]$**
-3.40 = log $[\text{H}^+]$
 $10^{-3.40} = 10^{\log[\text{H}^+]}$
 $3.98 \times 10^{-4} \text{ M} = [\text{H}^+]$
 $K_w = 1 \times 10^{-14} = [\text{H}^+][\text{OH}^-]$
 $[\text{OH}^-] = \frac{1 \times 10^{-14}}{[\text{H}^+]}$
 $[\text{OH}^-] = \frac{1 \times 10^{-14}}{3.98 \times 10^{-4} \text{ M}}$
 $[\text{OH}^-] = 2.51 \times 10^{-11} \text{ M}$
b) **pH = 6.7** (pH of saliva)
 $[\text{H}^+] = 1.99 \times 10^{-7} \text{ M}$
 $[\text{OH}^-] = 5.01 \times 10^{-8} \text{ M}$
d) **pOH = 2.15**
 $[\text{H}^+] = 1.41 \times 10^{-12} \text{ M}$
 $[\text{OH}^-] = 7.07 \times 10^{-3} \text{ M}$
e) **pOH = 12.4**
 $[\text{H}^+] = 2.51 \times 10^{-2} \text{ M}$
f) **pH = -0.650**
 $[\text{H}^+] = 4.47 \text{ M}$
 $[\text{OH}^-] = 2.24 \times 10^{-15} \text{ M}$

PS11.4. For each of the following acids, write the formula for the conjugate base.

- | | | |
|--------------------------------------|---|-----------------------------------|
| a) HPO_4^{2-} | c) H_2O | e) OH^- |
| PO_4^{3-} | OH^- | O^{2-} |
| b) HClO_3 | d) $\text{CH}_3\text{CH}_2\text{NH}_3^+$ | f) NH_4^+ |
| ClO_3^- | $\text{CH}_3\text{CH}_2\text{NH}_2$ | NH_3 |

PS11.5. For each of the following bases, write the formula for the conjugate acid.

- | | | |
|--|---|--|
| a) OH^- | c) HCO_3^{2-} | e) CH_3NH_2 |
| H_2O | H_2CO_3 | CH_3NH_3^+ |
| b) Cl^- | d) H_2O | f) $(\text{CH}_3)_3\text{N}$ |
| HCl | H_3O^+ | $(\text{CH}_3)_3\text{NH}^+$ |

PS11.6. For the following compounds, write the reaction with water and indicate the Brønsted acid, base, the conjugate acid and conjugate base.

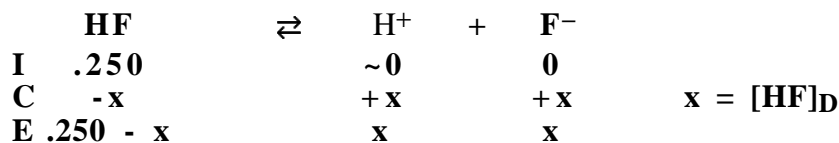
- | |
|--|
| a) $\text{HBr}(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{Br}^-(aq)$ |
| A B CA CB |
| b) $\text{NH}_3(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)$ |
| B A CA CB |
| c) $\text{HCN}(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CN}^-(aq)$ |
| A B CA CB |
| d) $\text{HC}_7\text{H}_5\text{O}_2(s) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{C}_7\text{H}_5\text{O}_2^-(aq)$ |
| A B CA CB |
| e) $\text{CH}_3\text{NH}_2(l) + \text{H}_2\text{O}(l) \rightleftharpoons \text{CH}_3\text{NH}_3^+(aq) + \text{OH}^-(aq)$ |
| B A CA CB |

PS11.7. For each of the following compounds, write two Brønsted-Lowry equations, one showing how the substance behaves as an acid, the second showing how the substance behaves as a base.

- | | |
|--|-------------|
| a) $\text{HCO}_3^-(aq)$ | |
| $\text{HCO}_3^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CO}_3^{2-}(aq)$ | acid |
| $\text{HCO}_3^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{OH}^-(aq) + \text{H}_2\text{CO}_3(aq)$ | base |
| b) $\text{NH}_3(aq)$ | |
| $\text{NH}_3(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)$ | base |
| $\text{NH}_3(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_2^-(aq) + \text{H}_3\text{O}^+(aq)$ | acid |
| c) $\text{HPO}_4^{2-}(aq)$ | |
| $\text{HPO}_4^{2-}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{PO}_4^{3-}(aq)$ | acid |
| $\text{HPO}_4^{2-}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{OH}^-(aq) + \text{H}_2\text{PO}_4^-(aq)$ | base |
| d) $\text{HSO}_4^-(aq)$ | |
| $\text{HSO}_4^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{SO}_4^{2-}(aq)$ | acid |
| $\text{HSO}_4^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{OH}^-(aq) + \text{H}_2\text{SO}_4(aq)$ | base |

PS11.8. Determine the equilibrium constant for the following solutions. (Show your work clearly!)

a) 0.250 M HF whose pH = 1.89.



$$K_a = \frac{[\text{H}^+][\text{F}^-]}{[\text{HF}]}$$

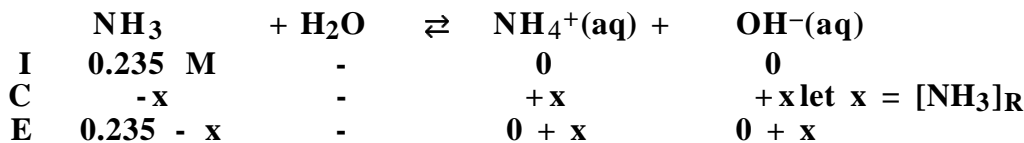
Since the pH = 1.89, the $[\text{H}^+] = 1.29 \times 10^{-2}$ M which is equal to x as shown in the ICE table above.

$$K_a = \frac{(x)(x)}{(0.250 - x)}$$

Substituting for x,

$$K_a = \frac{(1.29 \times 10^{-2})^2}{.250 - 0.0129} = 7.02 \times 10^{-4}$$

b) 0.235 M NH₃ whose pH = 11.31.

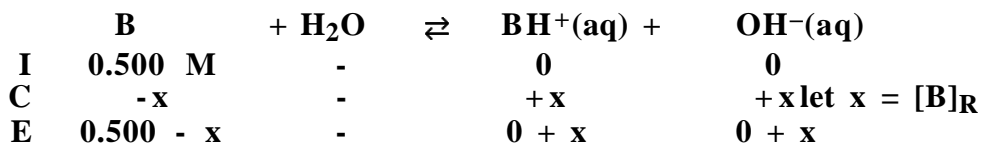


$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

Since the pH = 11.31, the pOH = 2.69 and the $[\text{OH}^-] = 2.0 \times 10^{-3}$ M which is also equal to x as shown in the ICE table above.

$$K_b = \frac{(x)(x)}{(0.1 - x)} = \frac{(2.0 \times 10^{-3})(2.0 \times 10^{-3})}{(0.235 - 2.0 \times 10^{-3})} = 1.79 \times 10^{-5}$$

c) 0.500 M B whose pH = 9.34.



$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

Since the pH = 9.34, the pOH = 4.66 and the $[\text{OH}^-] = 2.19 \times 10^{-5}$ M which is also equal to x as shown in the ICE table above.

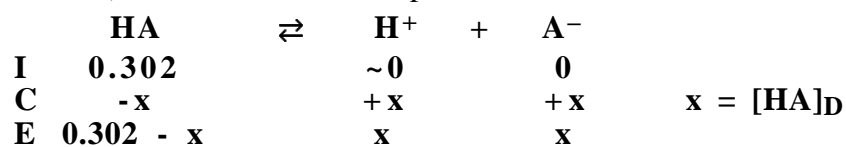
$$K_b = \frac{(x)(x)}{(0.5 - x)}$$

Substituting for x,

$$= \frac{(2.19 \times 10^{-5})(2.19 \times 10^{-5})}{(0.500 - 2.19 \times 10^{-5})}$$

$$K_b = \frac{(2.19 \times 10^{-5})^2}{0.500} = 9.57 \times 10^{-10}$$

d) 0.302 M HA whose pH = 4.80.



$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Since the pH = 4.80, the $[H^+] = 1.58 \times 10^{-5}$ M which is also equal to x as shown in the ICE table above.

$$K_a = \frac{(x)(x)}{(0.302 - x)}$$

Substituting for x,

$$K_a = \frac{(1.58 \times 10^{-5})^2}{.302 - 1.58 \times 10^{-5}}$$

$$= \frac{(1.58 \times 10^{-5})^2}{.302} = 8.32 \times 10^{-10}$$

PS11.9. Given the following substances and their initial concentration:

- | | | |
|--|--|---|
| a) 0.200 M HNO ₃ | e) 55.5 M H ₂ O | i) 0.200 M HC ₆ H ₅ O |
| b) 0.200 M HF | f) 0.200 M HNO ₂ | j) 0.200 M Ba(OH) ₂ |
| c) 0.200 M NaOH | g) 0.200 M CH ₃ NH ₂ | k) 0.003501 M HF |
| d) 0.200 M C ₅ H ₅ N | h) 0.200 M C ₂ H ₅ NH ₂ | l) 0.200 M HOCl |

Answer the following,

i) identify each as an acid, base or neutral substance.

- | | | | | | |
|--|----------|--|----------|---|----------|
| a) 0.200 M HNO ₃ | A | e) 55.5 M H ₂ O | N | i) 0.200 M HC ₆ H ₅ O | A |
| b) 0.200 M HF | A | f) 0.200 M HNO ₂ | A | j) 0.200 M Ba(OH) ₂ | B |
| c) 0.200 M NaOH | B | g) 0.200 M CH ₃ NH ₂ | B | k) 0.003501 M HF | A |
| d) 0.200 M C ₅ H ₅ N | B | h) 0.200 M C ₂ H ₅ NH ₂ | B | l) 0.200 M HOCl | A |

ii) list the K_a value for each acid and K_b value for each base.

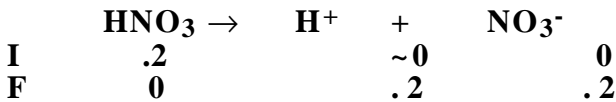
- | | |
|--|---|
| a) 0.200 M HNO ₃ | K_a is very large |
| b) 0.200 M HF | K_a = 6.8 x 10⁻⁴ |
| c) 0.200 M NaOH | K_b is very large |
| d) 0.200 M C ₅ H ₅ N | K_b = 1.7 x 10⁻⁹ |
| e) 55.5 M H ₂ O | K_a = 1 x 10⁻¹⁴ |
| f) 0.200 M HNO ₂ | K_a = 4.5 x 10⁻⁴ |
| g) 0.200 M CH ₃ NH ₂ | K_b = 4.4 x 10⁻⁴ |
| h) 0.200 M C ₂ H ₅ NH ₂ | K_b = 6.4 x 10⁻⁴ |
| i) 0.200 M HC ₆ H ₅ O | K_a = 1.3 x 10⁻¹⁰ |
| j) 0.200 M Ba(OH) ₂ | K_b is very large |
| k) 0.00491 M HF | K_a = 6.8 x 10⁻⁴ |
| l) 0.200 M HOCl | K_b = 3.0 x 10⁻⁸ |

iii) identify each substance as strong (S) or weak (W).

- | | | | | | |
|--|-----------|--|-----------|---|-----------|
| a) 0.200 M HNO ₃ | SA | e) 55.5 M H ₂ O | N | i) 0.200 M HC ₆ H ₅ O | WA |
| b) 0.200 M HF | WA | f) 0.200 M HNO ₂ | WA | j) 0.200 M Ba(OH) ₂ | SB |
| c) 0.200 M NaOH | SB | g) 0.200 M CH ₃ NH ₂ | WB | k) 0.003501 M HF | WA |
| d) 0.200 M C ₅ H ₅ N | WB | h) 0.200 M C ₂ H ₅ NH ₂ | WB | l) 0.200 M HOCl | WA |

iv) calculate the [H⁺] and the pH of each of the solutions.

- a) 0.200 M HNO₃

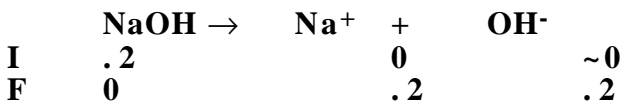


HNO₃ is a strong acid and completely dissociates in water.

Therefore, [H⁺] = 0.2 M and pH = 0.700

- b) 0.200 M HF [H⁺] = 1.2 x 10⁻² M pH = 1.93

- c) 0.100 M NaOH

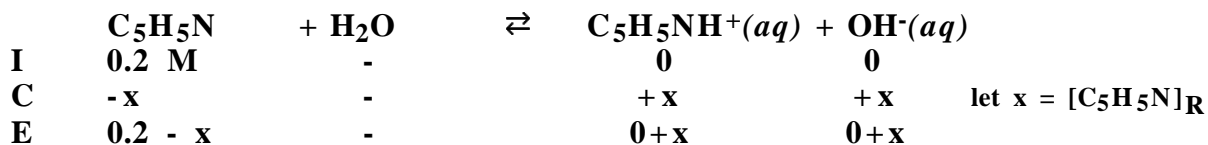


NaOH is a strong base and completely dissociates in water.

Therefore,

[OH⁻] = 0.2 M and pOH = 0.70, pH = 13.3: [H⁺] = 5.0 x 10⁻¹⁴ M

d) 0.200 M C₅H₅N



$$K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]}$$

$$1.7 \times 10^{-9} = \frac{(x)(x)}{0.2 - x} \quad \text{assume } x \ll 0.2$$

$$1.7 \times 10^{-9} = \frac{x^2}{.2}$$

$$3.4 \times 10^{-10} = x^2$$

$$1.8 \times 10^{-5} \text{ M} = x = [OH^-]$$

$$pOH = 4.73$$

$$pH = 9.27: [H^+] = 5.4 \times 10^{-10} \text{ M}$$

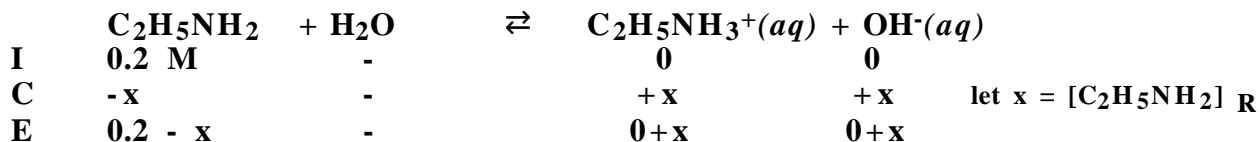
e) 55.5 M H₂O [H⁺] = 1.0 x 10⁻⁷ M: pH = 7.00

f) 0.200 M HNO₂ [H⁺] = 9.5 x 10⁻³ M: pH = 2.02

g) 0.200 M CH₃NH₂ [OH⁻] = 9.4 x 10⁻³ M and pOH = 2.03, pH = 11.98:

$$[H^+] = 1.07 \times 10^{-12} \text{ M}$$

h) 0.200 M C₂H₅NH₂



$$K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]}$$

$$6.4 \times 10^{-4} = \frac{(x)(x)}{0.2 - x} \quad \text{assume } x \ll 0.1$$

$$6.4 \times 10^{-4} = \frac{x^2}{.2}$$

$$1.28 \times 10^{-4} = x^2$$

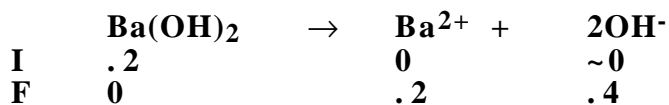
$$1.1 \times 10^{-2} \text{ M} = x = [OH^-]$$

$$pOH = 1.96$$

$$pH = 12.04: [H^+] = 9.09 \times 10^{-13}$$

i) 0.200 M HC₆H₅O [H⁺] = 5.1 x 10⁻⁶ M: pH = 5.30

j) 0.100 M Ba(OH)₂

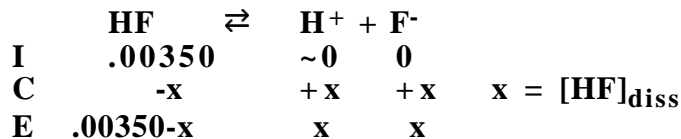


Ba(OH)₂ is a strong base and completely dissociates in water.

Therefore,

$$[OH^-] = 0.4 \text{ M and } pOH = 0.40, pH = 13.6: [H^+] = 2.5 \times 10^{-14} \text{ M}$$

k) 0.00350 M HF



$$K_a = \frac{[H^+][F^-]}{[HF]}$$

$$6.8 \times 10^{-4} = \frac{(x)(x)}{.00350-x}$$

$$6.8 \times 10^{-4} = \frac{x^2}{.00350-x}$$

$$2.38 \times 10^{-6} - 6.8 \times 10^{-4}x = x^2$$

$$2.38 \times 10^{-6} - 6.8 \times 10^{-4}x - x^2 = 0$$

$$x^2 + 6.8 \times 10^{-4}x - 2.38 \times 10^{-6} = 0$$

solving the quadratic equation $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$$x = \frac{-6.8 \times 10^{-4} \pm \sqrt{(6.8 \times 10^{-4})^2 - 4(1)(-2.38 \times 10^{-6})}}{2(1)}$$

$$x = \frac{-6.8 \times 10^{-4} \pm 3.16 \times 10^{-3}}{2}$$

Use only the positive root

$$x = 1.24 \times 10^{-3} \text{ M} = [H^+]$$

$$\text{pH} = 2.91$$

l) 0.200 M HOCl $[H^+] = 7.5 \times 10^{-5} \text{ M}$: $\text{pH} = 4.11$

v) determine the percent ionization for each acid and base.

w)

- | | |
|--|----------------|
| a) 0.200 M HNO ₃ | 100% |
| b) 0.200 M HF | 5.9% |
| c) 0.200 M NaOH | 100% |
| d) 0.200 M C ₅ H ₅ N | < 1% |
| e) 55.5 M H ₂ O | < 1% |
| f) 0.200 M HNO ₂ | 4.8% |
| g) 0.200 M CH ₃ NH ₂ | 4.7% |
| h) 0.200 M C ₂ H ₅ NH ₂ | 5.5% |
| i) 0.200 M HC ₆ H ₅ O | < 1% |
| j) 0.200 M Ba(OH) ₂ | 100% |
| k) 0.003501 M HF | 35.1% |
| l) 0.200 M HOCl | < 1% |

vi) rank all substances from strongest acid...weakest acid...neutrals.. ...weakest base...strongest base.

a)	0.200 M HNO ₃	pH = 0.700
b)	0.200 M HF	pH = 1.93
f)	0.200 M HNO ₂	pH = 2.02
l)	0.200 M HOCl	pH = 4.11
i)	0.200 M HC ₆ H ₅ O	pH = 5.30
e)	55.5 M H ₂ O	pH = 7.00
d)	0.200 M C ₅ H ₅ N	pH = 9.27
g)	0.200 M CH ₃ NH ₂	pH = 11.98
h)	0.200 M C ₂ H ₅ NH ₂	pH = 12.04
c)	0.200 M NaOH	pH = 13.3
j)	0.200 M Ba(OH) ₂	pH = 13.6